Student Textbook pages 374–375

For full solutions to the practice problems, visit **www.albertachemistry.ca**, Online Learning Centre, Instructor Edition, Full Solutions.

1. –44.2 kJ
2. –507 kJ
3. –121.0 kJ
4. 30 kJ
5. 52.6 kJ/mol
6. –128.8 kJ/mol

Student Textbook page 378

Q1. (a) $\frac{1}{2}H_2(g) + \frac{1}{2}I_2(g) \rightarrow HI(g)\ \Delta_f H^\circ = 26.5$ kJ/mol

(b) $2Fe(s) + \frac{3}{2}O_2(g) \rightarrow Fe_2O_3(s)$

$\Delta_f H^\circ = -824.2$ kJ/mol

(c) $Ba(s) + \frac{1}{2}O_2(g) \rightarrow BaO(s)\ \Delta_f H^\circ = -548.0$ kJ/mol

(d) $\frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g) + 2O_2(g) \rightarrow HClO_4(\ell)$

$\Delta_f H^\circ = -40.6$ kJ/mol

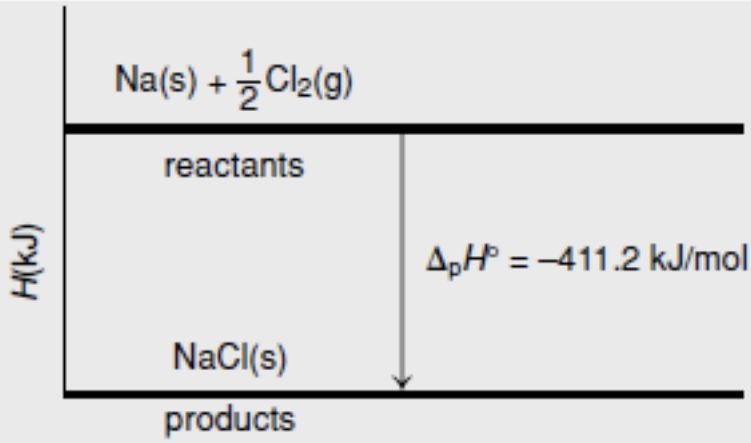
Q2. (a) endothermic

(b) exothermic

(c) exothermic

(d) exothermic

Q3.



Student Textbook page 379

Q4. (a) $C(s) + 2H_2(g) \rightarrow CH_4(g)$ $\Delta_f H^\circ = -74.6$ kJ/mol

(b) $6C(s) + 3H_2(g) \rightarrow C_6H_6(\ell)$ $\Delta_f H^\circ = 49.1$ kJ/mol

(c) $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$ $\Delta_f H^\circ = -411.2$ kJ/mol

(d) $Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$ $\Delta_f H^\circ = -601.6$ kJ/mol

(e) $Ca(s) + C(s) + \frac{3}{2}O_2(g) \rightarrow CaCO_3(s)$

$\Delta_f H^\circ = -1207.6$ kJ/mol

Q5. most thermally stable $CaCO_3(s)$, $MgO(s)$, $NaCl(s)$, $CH_4(g)$, $C_6H_6(\ell)$ least stable

Q6. $H_2(g) + \frac{1}{8}S_8(s) + 2O_2(g) \rightarrow H_2SO_4(\ell)$

$\Delta_f H^\circ = -814.0$ kJ/mol

Q7. formation: $6C(s) + 6H_2(g) + 3O_2(g) \rightarrow C_6H_{12}O_6(s)$

$\Delta_f H^\circ = -1273.3$ kJ/mol

photosynthesis: $6CO_2(g) + 6H_2O(\ell) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$ $\Delta H = 2802.5$ kJ/mol

Although both result in the formation of glucose, the heat of formation is exothermic, whereas the photosynthesis reaction is endothermic. The formation reaction forms glucose from its elements, and photosynthesis forms glucose from water and carbon dioxide. In a formation reaction, there are no by-products. Photosynthesis produces oxygen as a by-product.

Student Textbook pages 382–383

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7. –136.4 kJ

8. –882.6 kJ

9. –77.6 kJ

10. (a) –637.9 kJ/mol

(b) 2.49×10^3 kJ

11. –102.7 kJ/mol

12. –224.2 kJ/mol

Student Textbook page 383

1. When a reaction occurs in the opposite direction, the sign of the ΔH must be reversed due to the law of conservation of energy. For example, when glucose is formed in photosynthesis,

$$6CO_2(g) + 6H_2O(\ell) \rightarrow C_6H_{12}O_6(s) + 6O_2(g) \qquad \Delta H = 2802.5 \text{ kJ/mol}$$

2802.5 kJ/mol of glucose is stored as energy in the chemical bonds. When cellular respiration occurs, the reverse of photosynthesis, the same amount of energy that was stored, 2802.5 kJ/mol, will be released into the surroundings. Since this reaction is exothermic, the sign will be the negative of that of photosynthesis.

2. Most stable: NaCl(s), $H_2S(g)$, $O_2(g)$, $C_2H_4(g)$, $C_2H_2(g)$ least stable

3. (a) $6C(s) + 3H_2(g) \rightarrow C_6H_6(\ell)$ $\Delta_f H^\circ = 49.1$ kJ/mol

(b) Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= 0.4173 \text{ mol } C_6H_6(\ell)$$

$$\Delta H = n\Delta_f H$$
$$= (0.4173 \text{ mol } C_6H_6(\ell))\left(\frac{49.1 \text{ kJ}}{\text{mol } C_6H_6(\ell)}\right)$$
$$= 20.5 \text{ kJ}$$

Method 2 Using Dimensional Analysis

$$x \text{ kJ} = \frac{49.1 \text{ kJ}}{\text{mol } C_6H_6(\ell)} \times \frac{1 \text{ mol } C_6H_6(\ell)}{78.12 \text{ g } C_6H_6(\ell)} \times 32.6 \text{ g } C_6H_6(\ell)$$
$$= 20.5 \text{ kJ}$$

4. $2 \times$ (1) $2H_3BO_3(aq) \rightarrow 2HBO_2(aq) + 2H_2O(\ell)$ $\Delta H^\circ = 2(-0.02 \text{ kJ})$

$-\frac{1}{2} \times$ (2) $2HBO_2(aq) \rightarrow \frac{1}{2}H_2B_4O_7(s) + \frac{1}{2}H_2O(\ell)$ $\Delta H^\circ = -\frac{1}{2}(-11.3 \text{ kJ})$

$\frac{1}{2} \times$ (3) $\frac{1}{2}H_2B_4O_7(s) \rightarrow B_2O_3(s) + \frac{1}{2}H_2O(\ell)$ $\Delta H^\circ = \frac{1}{2}(17.5 \text{ kJ})$

$2H_3BO_3(aq) \rightarrow B_2O_3(s) + 3H_2O(\ell)$ $\Delta H^\circ = 14.4 \text{ kJ}$

5. Although mathematically, due to the law of conservation of energy, using enthalpies of formation is a shortcut for manipulating and adding equations, students should understand that in most reactions, the reactants do not actually break down into their elements and then react to form products.

Student Textbook page 387

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13. 18%

14. 69%

15. 0.88 g

16. 38.7%

Student Textbook page 392

1. The energy associated with physical changes are smallest, and those associated with nuclear processes are largest. The energy required to boil 1 mol of water is 40.65 kJ, the molar enthalpy of combustion of coal is about 3900 kJ, while the fusion of 1 mol of uranium-235 releases about 2.1×10^{10} kJ.
2. Students should ask the following questions: What definition of efficiency are you using? What is your definition of "useful energy"? What factors are included in your definition of "energy used"?
3. A simple experiment involves heating water in a beaker. Students will need to determine the flow rate of the gas and the time the gas burns, to find the quantity of gas burned. They should compare ideal enthalpy change for the burning to the amount of heat actually absorbed to increase the temperature of the water, say, by 30 °C. An experiment like this would give the efficiency of a laboratory burner in performing the specific task of heating water in a beaker. The efficiency definition is based on the ratio between the actual energy used to heat the water as compared with the ideal quantity of heat released by burning the gas.

4. Under operating conditions in a furnace where the water formed is gaseous, the enthalpy of combustion of methane is $\Delta_c H^\circ = -802.5$ kJ/mol.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \quad \Delta_c H^\circ = -802.5 \text{ kJ/mol}$$

However, if the gaseous water produced is condensed, further heat is released. The molar enthalpy of condensation for water is –40.65 kJ/mol, so if the two moles of water are condensed, an extra 81 kJ of energy is released. This increases the efficiency of the furnace by 10%.

5. (a) The manufacturer is probably using a definition that compares electrical energy used by the device to thermal energy gained by the water in the kettle.

(b) $$\text{Efficiency} = \left| \frac{\text{Energy output}}{\text{Energy input}} \right| \times 100\%$$

$$= \left| \frac{\text{heat absorbed by water}}{\text{electrical energy used}} \right| \times 100\%$$

$$= 95\%$$

(c) Students should outline an experiment that monitors the electrical energy used to raise a measured volume of water in the kettle to 100 °C, comparing it to the calculated energy change based on the specific heat capacity of water. Safety precautions should include care using electricity and water.

6. From a purely environmental perspective, the "best" barbecue would be the natural gas barbecue. From a purely economical perspective, the choice would be a hibachi. The propane barbecue may be considered a compromise, however its lower cost may not compensate for the environmental cost of propane over natural gas. Students may also consider the availability of propane (if they don't have a natural gas connection available) or the convenience of natural gas if they do have a connection available.

7. (a) Part 1 Energy input

Since the methane burner is heating a beaker of water, we can assume it is an open system. From Table 9.3, the molar enthalpy of combustion of methane, $CH_4(g)$, is –802.5 kJ/mol.

Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= \frac{0.37 \text{ g } \cancel{CH_4(g)}}{\frac{16.05 \text{ g } \cancel{CH_4(g)}}{\text{mol } CH_4(g)}}$$

$$= 0.0231 \text{ mol } CH_4(g)$$

$$\Delta H = n\Delta H_c$$

$$= (0.0231 \text{ } \cancel{\text{mol } CH_4(g)}) \left(\frac{-802.5 \text{ kJ}}{\cancel{\text{mol } CH_4(g)}} \right)$$

$$= -18.5 \text{ kJ}$$

Method 2 Using Dimensional Analysis

$$\text{x kJ} = 0.37 \text{ } \cancel{\text{g } CH_4(g)} \times \frac{1 \text{ } \cancel{\text{mol } CH_4(g)}}{16.05 \text{ } \cancel{\text{g } CH_4(g)}} \times \frac{-802.5 \text{ kJ}}{\cancel{\text{mol } CH_4(g)}}$$

$$= -18.5 \text{ kJ}$$

Part 2 Energy output

Mass of water = 427.96 g – 275.38 g = 152.58 g

Temperature change = 37.9 °C – 25.3 °C = 12.6 °C

$$Q = (mc\Delta t)_{water} + (mc\Delta t)_{glass}$$

$$= (152.58\ \cancel{g})\left(\frac{4.19\ \text{J}}{\cancel{g}^{\bullet}\cancel{^\circ C}}\right)(12.6\ \cancel{^\circ C}) + (275.38\ \cancel{g})\left(\frac{0.84\ \text{J}}{\cancel{g}^{\bullet}\cancel{^\circ C}}\right)(12.6\ \cancel{^\circ C})$$

$$= 11.0\ \text{kJ}$$

$$\text{Efficiency} = \left|\frac{\text{Energy output}}{\text{Energy input}}\right| \times 100\%$$

$$= \left|\frac{11.0\ \text{kJ}}{-18.5\ \text{kJ}}\right| \times 100\%$$

$$= 60\%$$

(b) The majority of the wasted energy would have been used to heat up the surroundings, which includes the air surrounding the burner, the burner itself, the thermometer and stirring rod.

(c) To decrease heat lost and reduce wasted energy, the system must be insulated as best as possible from the surroundings. By shielding the beaker from the outside air with an insulating substance such as Styrofoam™, and by putting a lid on the beaker, this would minimize heat loss and therefore increase efficiency.

8. (a) This is worse than the posted fuel economy, since the gas consumption is higher than the posted consumptions.

(b) Some reasons may include: driving at higher speeds in the city, idling the engine, driving with under-inflated tires, accelerating quickly from a stop, not coasting to a stop (all would increase fuel consumption). As well, a new car with an engine that is not yet broken in will have higher fuel consumption.

(c) Fuel economy can be improved by changing your driving habits. Driving slower, not idling the car, accelerating slowly from a stop and allowing a car to slow down by taking your foot off the gas rather than applying the brakes will all improve fuel economy.

9. To use the waste heat produced by any cooling system, which could include an air-conditioner as well as a refrigerator, you could build something similar to a solar-heating system. The warm air exiting the air conditioner or refrigerator would have to be contained and circulated around piping containing water. The preheated water would then already be partially warmed before entering the hot water tank, reducing energy costs. Heated water could also be used as radiant floor heating.

Student Textbook pages 400–401

Answers to Understanding Concepts Questions

1. If the enthalpy change for the original equation is ΔH, then the enthalpy change for the manipulated equation will be $3 \times -\Delta H$. The molar enthalpy change of the reaction is not changed.

2. (a) (a) $Ag(s) + \frac{1}{2}Cl_2(g) \rightarrow AgCl(s)$

$\Delta_f H^\circ = -127.0$ kJ/mol

(b) $2C(s) + 3H_2(g) + \frac{1}{2}O_2(g) \rightarrow C_2H_5OH(\ell)$

$\Delta_f H^\circ = -277.6$ kJ/mol

(c) $N_2(g) + 2H_2(g) + \frac{3}{2}O_2(g) \rightarrow NH_4NO_3(s)$

$\Delta_f H^\circ = -365.6$ kJ/mol

(b) Least stable to most stable: $AgCl(s)$, $C_2H_5OH(\ell)$, $NH_4NO_3(s)$

3. The standard state of an element is its most stable form at SATP (25 °C and 100 kPa). At these conditions, iodine is a solid. For 1 mol of $I_2(s)$ to sublimate and form $I_2(g)$, 21 kJ of energy must be absorbed.

4. Efficiency is usually defined as useful energy output divided by the energy input, expressed as a percentage. In this case, efficiency would be best described as moving the most people using the least amount of energy. In this case, the bus would be more efficient than a car.
5. The lawn could be mowed using a more efficient electric mower, or an even more efficient push mower.
6. The useful energy output by a hair dryer would be the moving hot air (using energy to heat the air and move the air). The rest of the energy would be waste. This would include the noise made by the hair dryer and the heating of the hair dryer itself.
7. A smart shopper should look for an EnerGuide and an ENERGY STAR® sticker on their appliances. The EnerGuide label allows a consumer to quickly judge the energy efficiency of the appliance compared to others in its class using a sliding scale. If present, the ENERGY STAR® label is an easy way to tell that this appliance is one of the most energy efficient in its category.
8. A car will operate at its maximum fuel efficiency if it is well maintained and has its tires inflated to the proper pressure. By driving at or below the speed limit and by not accelerating or decelerating quickly, a driver can further improve their fuel efficiency.
9. Using hydro or wind to produce electricity is more efficient than using coal or natural gas because there are fewer energy conversions. Wind turns the turbines directly to produce electricity, whereas coal must be burned to heat water and create steam to turn turbines. There are more steps, so more energy is lost along the way.

Answers to Applying Concepts Questions

10. Step 1: Determine the heat of combustion of carbon in excess oxygen to form carbon dioxide:

(1) $C(s) + O_2(g) \rightarrow CO_2(g)\ \Delta H_1$

Step 2: Determine the enthalpy change for the combustion of CO(g) in excess oxygen to form $CO_2(g)$:

(2) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)\ \Delta H_2$

The heat of formation of CO(g) can be determined using these two equations and Hess's law:

$$CO_2(g) \rightarrow CO(g) + \frac{1}{2}O_2(g) \qquad -\Delta H_2$$

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H_1$$

$$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g) \qquad \Delta_f H^\circ = \Delta H_1 - \Delta H_2$$

11. There are many ways to reduce greenhouse gas emissions and still keep driving a car. Making some simple changes in driving habits and performing regular maintenance will make any car more fuel-efficient. As well, energy can be saved at home by having shorter showers and by using a low-flow showerhead. Students could direct their friend to one of the websites they have used during this chapter to investigate ways to conserve energy and reduce greenhouse gas emissions.

12. (a) Since the overall enthalpy change for the reaction is negative and the enthalpy of formation of hydrogen, $H_2(g)$, is zero, then the potential energy stored in acetylene, $C_2H_2(g)$ must be higher (have a higher enthalpy of formation) than ethylene, $C_2H_4(g)$.

(b) Acetylene would release more energy when burned since it has a higher enthalpy of formation; it has more stored energy in its three bonds, which would be released when burned.

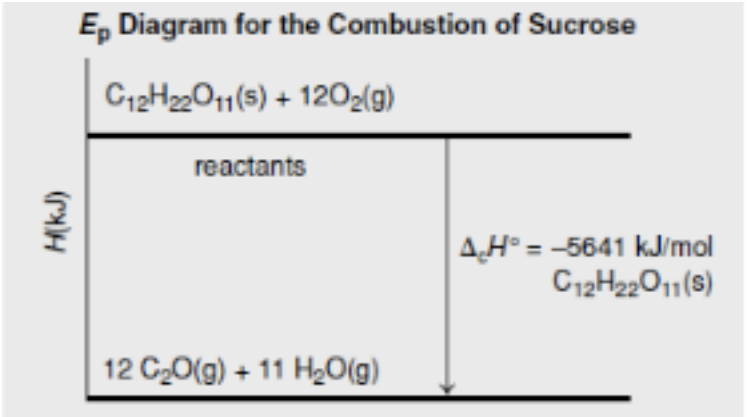
13. Students should disagree with this statement. Although harnessing electricity does not directly produce any emissions, building and running the plant does. In constructing a hydroelectric power plant, ecosystems are permanently destroyed, therefore, there is considerable harm to the environment.

14. A primary reason why nuclear energy is not a viable alternative for Alberta is that these plants require a significant amount of water for cooling. Although uranium is mined in Saskatchewan, it is processed in Ontario, which would add shipping costs to the cost of producing energy. Since Alberta is rich in coal and other fossil fuels, it is economically better to produce energy using closely available resources. To decide whether it would be feasible to produced energy using nuclear power, students should perform a risk-benefit analysis. Students could include the following topics: available technology, economic issues, availability of fossil fuels and uranium, potential hazards to the environment, effects on the Alberta economy, and potential health problems resulting from the use of these fuels.

Answers to Solving Problems Questions

15. (a) $-5641\ \text{kJ} = [(12\ \text{mol})(-393.5\ \text{kJ/mol}) + (11\ \text{mol})(-285.8\ \text{kJ/mol})] - [(12\ \text{mol})(0\ \text{kJ/mol}) + (1\ \text{mol})(\Delta_f H^\circ\ C_{12}H_{22}O_{11}(g))]$

$\Delta_f H^\circ\ C_{12}H_{22}O_{11}(g) = -2225\ \text{kJ/mol}$



16. $-1 \times (3)\ CHCl_3(\ell) \rightarrow \cancel{C(s)} + \frac{1}{2}\cancel{H_2(g)} + \frac{3}{2}Cl_2(g)$

$\Delta H^\circ = -1(-134.5\ kJ)$

$-3 \times (1)\ 3HCl(g) \rightarrow \frac{3}{2}\cancel{H_2(g)} + \frac{3}{2}Cl_2(g)$

$\Delta H^\circ = -3(-92.3\ kJ)$

$(2)\ \cancel{C(s)} + 2\cancel{H_2(g)} \rightarrow CH_4(g) \quad \Delta H^\circ = -74.6\ kJ$

$CHCl_3(\ell) + 3HCl(g) \rightarrow CH_4(g) + 3Cl_2(g)$

$\Delta H^\circ = (134.5\ kJ + 276.9\ kJ - 74.6\ kJ)$

$= 336.8\ kJ$

17. $\Delta H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants})$

$-1178\ kJ = [(2\ mol \times \Delta_f H^\circ\ CO_2(g)) + (3\ mol \times \Delta_f H^\circ\ H_2O(\ell))] - [(1\ mol \times \Delta_f H^\circ\ (CH_2OH)_2(\ell)) + 0]$

$-1178\ kJ = [(2\ mol \times -393.5\ kJ/mol) + (3\ mol \times -285.8\ kJ/mol)] - [1\ mol \times \Delta_f H^\circ (CH_2OH)_2(\ell)]$

$-1178\ kJ = -787.0\ kJ - 857.4\ kJ - (1\ mol) \times (\Delta_f H^\circ\ (CH_2OH)_2(\ell))$

$\Delta_f H^\circ\ (CH_2OH)_2(\ell) = -466\ kJ/mol$

18. (a) $H_2(g) + O_2(g) \rightarrow H_2O_2(\ell)$

(b) $-\frac{1}{2} \times (1)\ \cancel{H_2O}(\ell) + \frac{1}{2}O_2(g) \rightarrow H_2O_2(\ell)$

$\Delta H^\circ = -\frac{1}{2}(-196\ kJ)$

$(1)\ H_2(g) + \frac{1}{2}O_2(g) \rightarrow \cancel{H_2O}(\ell)$

$\Delta H^\circ = -286\ kJ$

$H_2(g) + O_2(g) \rightarrow H_2O_2(\ell)$

$\Delta H^\circ = (+98\ kJ - 286\ kJ)$

$\Delta H^\circ = -188\ kJ$

19. $\frac{1}{2} \times$ (1) $\frac{1}{2}H_2(g) + C(s) + \frac{1}{2}N_2(g) \rightarrow HCN(g)$

$$\Delta H^\circ = \frac{1}{2}(270 \text{ kJ})$$

$-1 \times$ (3) $CH_4(g) \rightarrow C(s) + 2H_2(g)$

$$\Delta H^\circ = -1(-75 \text{ kJ})$$

$-\frac{1}{2} \times$ (2) $NH_3(g) \rightarrow \frac{1}{2}N_2(g) + \frac{3}{2}H_2(g)$

$$\Delta H^\circ = -\frac{1}{2}(-92 \text{ kJ})$$

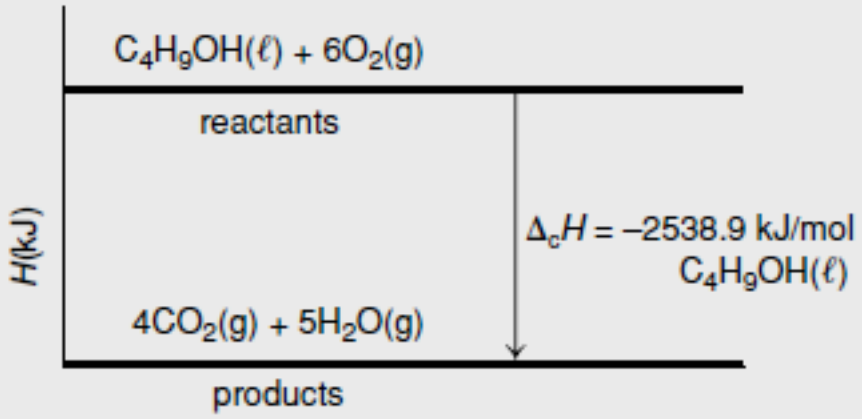
$$CH_4(g) + NH_3(g) \rightarrow HCN(g) + 3H_2(g)$$
$$\Delta H^\circ = (+135 \text{ kJ} + 75 \text{ kJ} + 46 \text{ kJ})$$
$$\Delta H^\circ = +256 \text{ kJ}$$

20. (a) $\Delta_c H = [4(\Delta_f H^\circ\ CO_2(g)) + 5(\Delta_f H^\circ\ H_2O(g))] - [6(\Delta_f H^\circ\ O_2(g)) + 1(\Delta_f H^\circ\ C_4H_9OH(\ell))]$

$\Delta_c H = [(4 \text{ mol})(-393.5 \text{ kJ/mol}) + (5 \text{ mol})(-241.8 \text{ kJ/mol})] - [(6 \text{ mol})(0 \text{ kJ/mol}) + (1 \text{ mol})(-244.4 \text{ kJ/mol})]$

$\Delta_c H = -2538.6 \text{ kJ/mol}$

(b)



(c) Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= \frac{10.0 \text{ g } C_4H_9OH(\ell)}{\frac{74.14 \text{ g } C_4H_9OH(\ell)}{\text{mol } C_4H_9OH(\ell)}}$$

$$= 0.1349 \text{ mol } C_4H_9OH(\ell)$$

$$\Delta H = n\Delta_c H$$

$$= (0.1349 \text{ mol } C_4H_9OH(\ell))\left(\frac{-2538.6 \text{ kJ}}{\text{mol } C_4H_9OH(\ell)}\right)$$

$$= -342 \text{ kJ}$$

Method 2 Using Dimensional Analysis

$$\text{x kJ} = 10.0 \text{ g } C_4H_9OH(\ell) \times \frac{1 \text{ mol } C_4H_9OH(\ell)}{74.14 \text{ g } C_4H_9OH(\ell)} \times \frac{-2538.6 \text{ kJ}}{\text{mol } C_4H_9OH(\ell)}$$

$$= -342 \text{ kJ}$$

21. First, the molar enthalpy of combustion of ethanol must be calculated:

$$C_2H_5OH(\ell) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

$$\Delta H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants})$$

$$= [(2 \text{ mol})(\Delta_f H^\circ CO_2(g)) + (3 \text{ mol})(\Delta_f H^\circ H_2O(g))] - [(1 \text{ mol})(\Delta_f H^\circ C_2H_5OH(\ell)) + (3 \text{ mol})(\Delta_f H^\circ O_2(g))]$$

$$= [(2 \text{ mol})(-393.5 \text{ kJ/mol}) + (3 \text{ mol})(-241.8 \text{ kJ/mol})] - [(1 \text{ mol})(-277.6 \text{ kJ/mol}) + (3 \text{ mol})(0 \text{ kJ/mol})]$$

$$= -1512.4 \text{ kJ} - (-277.6 \text{ kJ})$$

$$= -1234.8 \text{ kJ/mol } C_2H_5OH(\ell)$$

The molar enthalpy of combustion from question 20 is –2538.6 kJ/mol.

The amount of energy released by the mixture can be calculated by determining the amount of energy released by each of the components and then adding them together. In a 1.0 g sample, there would be 0.70 g of ethanol and 0.30 g of butan-1-ol. By determining how much energy is released by these quantities of each fuel, a total amount of energy per gram of the fuel in the burner can be determined.

Method 1 Using Formulas

For butan-1-ol

$$n = \frac{m}{M}$$

$$= \frac{0.30 \text{ g } \cancel{C_4H_9OH(\ell)}}{\frac{74.14 \text{ g } \cancel{C_4H_9OH(\ell)}}{\text{mol } C_4H_9OH(\ell)}}$$

$$= 0.00405 \text{ mol } C_4H_9OH(\ell)$$

$$\Delta H = n\Delta_c H$$

$$= (0.00405 \text{ } \cancel{\text{mol } C_4H_9OH(\ell)}) \left(\frac{-2538.6 \text{ kJ}}{\cancel{\text{mol } C_4H_9OH(\ell)}} \right)$$

$$= -10 \text{ kJ}$$

For ethanol

$$n = \frac{m}{M}$$

$$= \frac{0.70 \text{ g } \cancel{C_2H_5OH(\ell)}}{\frac{46.08 \text{ g } \cancel{C_2H_5OH(\ell)}}{\text{mol } C_2H_5OH(\ell)}}$$

$$= 0.0152 \text{ mol } C_2H_5OH(\ell)$$

$$\Delta H = n\Delta_c H$$
$$= (0.0152 \text{ mol } C_2H_5OH(\ell))\left(\frac{-1234.8 \text{ kJ}}{\text{mol } C_2H_5OH(\ell)}\right)$$
$$= -19 \text{ kJ}$$

Therefore, the total amount of energy released per gram of fuel can be determined:

$$\Delta H = -10 \text{ kJ} + (-19 \text{ kJ})$$
$$= -29 \text{ kJ/g of burner fuel}$$

Method 2 Using Dimensional Analysis

$$\text{x kJ} = 0.30 \text{ g } C_4H_9OH(\ell) \times \frac{1 \text{ mol } C_4H_9OH(\ell)}{74.14 \text{ g } C_4H_9OH(\ell)} \times \frac{-2538.6 \text{ kJ}}{\text{mol } C_4H_9OH(\ell)} + 0.70 \text{ g } C_2H_5OH(\ell) \times \frac{1 \text{ mol } C_2H_5OH(\ell)}{46.08 \text{ g } C_2H_5OH(\ell)} \times \frac{-1234.8 \text{ kJ}}{\text{mol } C_2H_5OH(\ell)}$$

$$= -29 \text{ kJ/g of burner fuel}$$

22. First, the enthalpy change for the reaction must be calculated:

$$3H_2(g) + CO(g) \rightarrow CH_4(g) + H_2O(g)$$
$$\Delta H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants})$$
$$= [(1 \text{ mol})(\Delta_f H^\circ CH_4(g)) + (1 \text{ mol})(\Delta_f H^\circ H_2O(g))] - [(3 \text{ mol})(\Delta_f H^\circ H_2(g)) + (1 \text{ mol})(\Delta_f H^\circ CO(g))]$$
$$= [(1 \text{ mol})(-74.6 \text{ kJ/mol}) + (1 \text{ mol})(-241.8 \text{ kJ/mol})] - [(3 \text{ mol})(0) + (1 \text{ mol})(-110.5 \text{ kJ/mol})]$$
$$= -205.9 \text{ kJ}$$

Now, you can find the quantity of energy released or absorbed:

Method 1 Using Formulas

$$n = \frac{m}{M}$$

$$= \frac{7.9 \text{ g } \cancel{CH_4(g)}}{\frac{16.05 \text{ g } \cancel{CH_4(g)}}{\text{mol } CH_4(g)}}$$

$$= 0.492 \text{ mol } CH_4(g)$$

$$\Delta H = n\Delta_r H$$
$$= (0.492 \text{ } \cancel{\text{mol } CH_4(g)}) \left(\frac{-205.9 \text{ kJ}}{\cancel{\text{mol } CH_4(g)}} \right)$$
$$= -1.0 \times 10^2 \text{ kJ}$$

Method 2 Using Dimensional Analysis

$$\text{x kJ} = 7.9 \text{ g } \cancel{CH_4(g)} \times \frac{1 \text{ } \cancel{\text{mol } CH_4(g)}}{16.05 \text{ g } \cancel{CH_4(g)}} \times \frac{-205.9 \text{ kJ}}{\cancel{\text{mol } CH_4(g)}}$$
$$= -1.0 \times 10^2 \text{ kJ}$$

23. (a), (b), (c) Accept all reasonable procedures. Students will probably design an experiment similar to Investigation 9.B on page 358-359, using a measured mass of camping fuel in an alcohol burner (or other similar burner) to heat water in a calorimeter. They can determine the enthalpy change based on the temperature change of the water.

(d) The experimental value is likely lower, as the accepted value for the molar enthalpy of combustion assumes there are no heat losses to the surroundings, and the heat content of the calorimeter is negligible. The accepted value is also based on an assumption that complete combustion takes place, although this is probably not the case.

24. Energy input = mass of fuel × enthalpy of combustion

$$= (3.25\text{ g})(-44.6\text{ kJ/g})$$
$$= -145\text{ kJ}$$

Energy Ouput

$$Q = mc\Delta t_{\text{water}} + mc\Delta t_{\text{aluminium}}$$
$$= (250\text{ g})\left(\frac{4.19\text{ J}}{\text{g}\cdot{}^\circ\text{C}}\right)(85.0\,{}^\circ\text{C} - 2.5\,{}^\circ\text{C}) + (150\text{ g})\left(\frac{0.897\text{ J}}{\text{g}\cdot{}^\circ\text{C}}\right)(85.0\,{}^\circ\text{C} - 2.5\,{}^\circ\text{C})$$
$$= 97.5\text{ kJ}$$

$$\text{Efficiency} = \left|\frac{\text{Energy output}}{\text{Energy input}}\right| \times 100\%$$
$$= \left|\frac{97.5}{-145}\right| \times 100\%$$
$$= 67\%$$

25. Energy input is equal to the amount of energy provided by 4.75 g of methane. The molar enthalpy of combination of methane in an open system can be found in Table 9.1 on page 347.

Energy input

Method 1 Using Formulas

$$n = \frac{m}{M}$$
$$= \frac{4.75\text{ g } \cancel{\text{CH}_4\text{(g)}}}{\dfrac{16.05\text{ g } \cancel{\text{CH}_4\text{(g)}}}{\text{mol CH}_4\text{(g)}}}$$
$$= 0.2960\text{ mol CH}_4\text{(g)}$$

$$\Delta H = n\Delta_c H$$
$$= (0.2960\ \cancel{\text{mol CH}_4\text{(g)}})\left(\frac{-802.5\text{ kJ}}{\cancel{\text{mol CH}_4\text{(g)}}}\right)$$
$$= -238\text{ kJ}$$

Method 2 Using Dimensional Analysis

$$x\ \text{kJ} = 4.75\ \text{g}\ \cancel{CH_4(g)} \times \frac{1\ \cancel{\text{mol}\ CH_4(g)}}{16.05\ \cancel{\text{g}\ CH_4(g)}} \times \frac{-802.5\ \text{kJ}}{\cancel{\text{mol}\ CH_4(g)}}$$

$$= -238\ \text{kJ}$$

$$\text{Efficiency} = \left| \frac{\text{Energy output}}{\text{Energy input}} \right| \times 100\%$$

$$47\ \% = \left| \frac{\text{Energy output}}{-238\ \text{kJ}} \right| \times 100\%$$

Energy output = 112 kJ

Use the energy output to determine the mass of water that could be heated.

$$Q = mc\Delta t$$

$$m = \frac{Q}{c\Delta t}$$

$$= \frac{1.12 \times 10^5\ \cancel{J}}{\left(\frac{4.19\ \cancel{J}}{\text{g}\cdot\cancel{^\circ\text{C}}}\right)(100\ \cancel{^\circ\text{C}} - 10\ \cancel{^\circ\text{C}})}$$

$$= 3.0 \times 10^2\ \text{g}$$

26. First, calculate the amount of energy required to preheat the barbeque (energy output)

$$Q = mc\Delta t$$

$$= (5000\ \cancel{\text{g}})\left(\frac{0.46\ \text{J}}{\cancel{\text{g}}\cdot{}^\circ\text{C}}\right)(175\ ^\circ\text{C} - 25\ ^\circ\text{C})$$

$$= 345\ \text{kJ}$$

$$\text{Efficiency} = \left| \frac{\text{Energy output}}{\text{Energy input}} \right| \times 100\%$$

$$50\% = \left| \frac{345\ \text{kJ}}{\text{Energy input}} \right| \times 100\%$$

Energy input = 690 kJ

Use the required energy input and the molar enthalpy of combustion of propane (found in Table 9.1 on page 347) to determine the mass of propane required.

Method 1 Using Formulas

$$\Delta H = n\Delta_c H$$

$$n = \frac{\Delta H}{\Delta_c H}$$

$$= \frac{-690\ \cancel{kJ}}{\frac{-2043.9\ \cancel{kJ}}{\text{mol } C_3H_8(g)}}$$

$$= 0.338 \text{ mol } C_3H_8(g)$$

$$n = \frac{m}{M}$$

$$m = nM$$

$$= (0.338\ \cancel{\text{mol } C_3H_8(g)})\left(\frac{44.11 \text{ g } C_3H_8(g)}{\cancel{\text{mol } C_3H_8(g)}}\right)$$

$$= 15 \text{ g } C_3H_8(g)$$

Method 2 Using Dimensional Analysis

$$\text{x g } C_3H_8(g) = -690\ \cancel{kJ} \times \frac{\cancel{\text{mol } C_3H_8(g)}}{-2043.9 \text{ kJ}} \times \frac{44.11 \text{ g } C_3H_8(g)}{\cancel{\text{mol } C_3H_8(g)}}$$

$$= 15 \text{ g } C_3H_8(g)$$

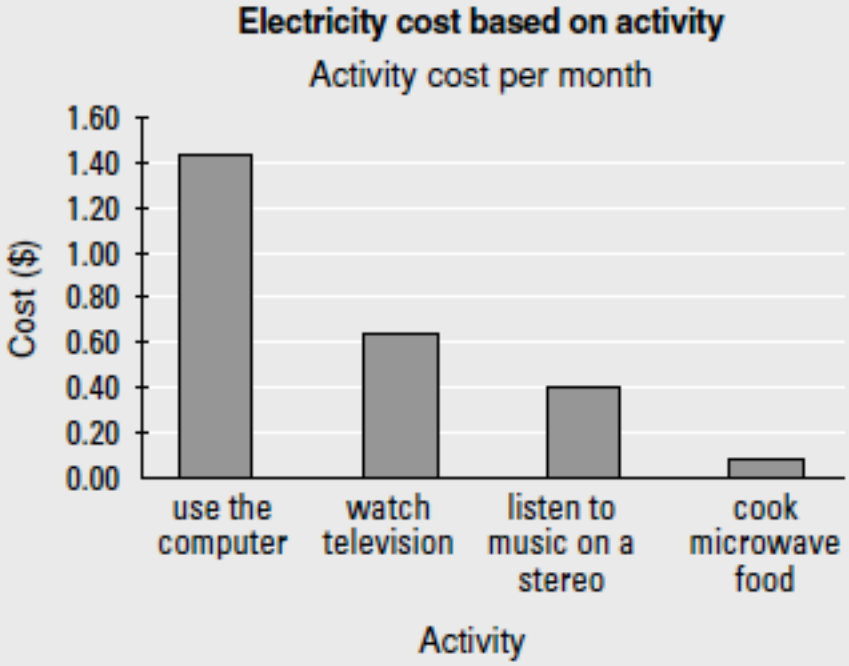
Answers to Making Connections Questions

27. (a) and **(b)** Sample calculations are shown below:

Based on an electrical cost of $ 0.06/kWh

Activity	Typical Wattage (W)	Hours/ month used (h)	Cost per month ($)
use the computer	200	120	1.44
watch television	180	60	0.72
listen to music on a stereo	115	60	0.41
cook microwave food	600	2.5	0.09

(c)



(d) The activities that use the most energy, and therefore cost the most, are likely ones that students will not think of, and include heating their homes, the water they shower in and the laundry they wash. Of the activities that students will likely list, the computer is probably the most energy-consuming.

(e) Student spreadsheets should include a table such as the one above, where the electrical cost would be entered into a cell and then referenced in the calculation for "Cost per month".

28. (a) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

$\Delta_c H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants})$

$= [(1 \text{ mol})(\Delta_f H^\circ CO_2(g)) + (2 \text{ mol})(\Delta_f H^\circ H_2O(g))] - [(1 \text{ mol})(\Delta_f H^\circ CH_4(g)) + (2 \text{ mol})(\Delta_f H^\circ O_2(g))]$

$= [(1 \text{ mol})(-393.5 \text{ kJ/mol}) + (2 \text{ mol})(-241.8 \text{ kJ/mol})] - [(1 \text{ mol})(-74.6 \text{ kJ/mol}) + (2 \text{ mol})(0 \text{ kJ/mol})]$

$= -802.5 \text{ kJ/mol } CH_4(g)$

$$x\ \frac{\text{kJ}}{\text{g } CH_4(g)} = \frac{-802.5\ \text{kJ}}{\cancel{\text{mol } CH_4(g)}} \times \frac{\cancel{\text{mol } CH_4(g)}}{16.05\ \text{g } CH_4(g)}$$

$$= -50.0\ \frac{\text{kJ}}{\text{g } CH_4(g)}$$

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)$$

$$\Delta_c H^\circ = \Delta_f H^\circ\ H_2O(g) = -241.8\ \text{kJ/mol } H_2(g)$$

$$\times\ \frac{\text{kJ}}{\text{g } H_2(g)} = \frac{-241.8\ \text{kJ}}{\cancel{\text{mol } H_2(g)}} \times \frac{\cancel{\text{mol } H_2(g)}}{2.02\ \text{g } H_2(g)}$$

$$= -120\ \frac{\text{kJ}}{\text{g } H_2(g)}$$

(b) Hydrogen, $H_2(g)$, has a higher heat content.

(c) If both fuels burned completely, then hydrogen would allow you to drive farther since it has a higher heat content.

(d) Methane (natural gas), a fossil fuel, can be drilled for and found along with crude oil, or by itself, in pockets located in sedimentary rock. It can also be collected from decomposing material in landfills and sewage treatment plants. Hydrogen is most commonly obtained from the electrolysis of water. Currently, producing methane is much less expensive than producing hydrogen as the technology is in place and in use on a large scale.

(e) Students will likely consider hydrogen the more environmentally-friendly fuel, since the combustion of hydrogen produces only water and energy. The production of hydrogen, however, consumes electricity, which has its own environmental costs. There is an environmental cost to the production of methane and its combustion (primarily greenhouse gases), however, unlike hydrogen, the energy is already stored in the methane, so there is no environmental cost in its "production."